

Name: \_\_\_\_\_

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## VOLTAIC AND ELECTROLYTIC CELLS

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### Overview:

**Part A: Determining the Voltage of Electrochemical Cells**

**Part B: Examination of Electrolytic cell using carbon electrodes in a KI solution**

**Part C: Electroplating Pennies and calculating Faraday's Law**

### Part A: Determining the Voltage of Voltaic Cells

**Read and/or review Chapter 19: Sections 2-5, 7, 10, and 11 in your textbook.**

*Voltaic (Galvanic) Cells are electrochemical cells in which a spontaneous reaction generates an electric current.* Voltaic cells have two half-cells connected, so that electrons flow from one metal electrode to another through an external circuit and the ions flow through an internal cell connection (or salt bridge). The half-cell in which a half-reaction occurs with a loss of electrons (oxidation) is the **anode**. The second half-cell, in which a half-reaction occurs with a gain of electrons (reduction), is the **cathode**. Oxidation always occurs at the **anode** and reduction always occurs at the **cathode**. A half-reaction can also be referred to as a **redox couple**.

The objective of the first part of the lab is to determine the cell potential ( $E_{\text{cell}}$ ) for various voltaic cells and compare the data with the calculated  $E_{\text{cell}}$  values obtained by using the Nernst equation.

$$E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$$

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - (0.0592/n) \log Q$$

$n = \# \text{ of electrons}$

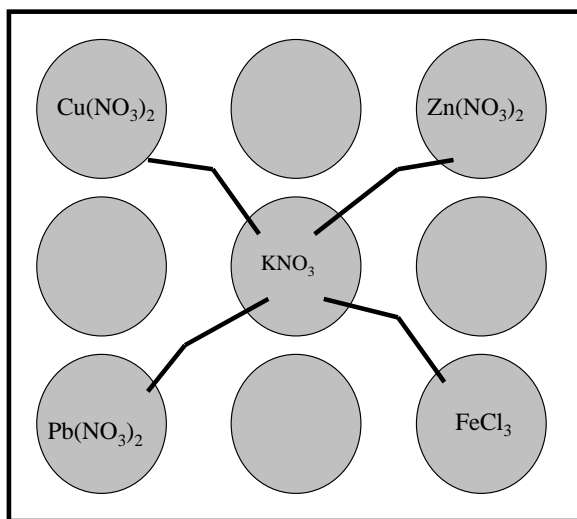
$Q = \text{reaction quotient}$

### Procedure:

1. Place about 2 mL of solutions 0.1 M  $\text{Cu}(\text{NO}_3)_2$ , 0.1 M  $\text{Zn}(\text{NO}_3)_2$ , 0.1 M  $\text{Pb}(\text{NO}_3)_2$ , 0.1 M  $\text{FeCl}_3$  and 0.1 M  $\text{KNO}_3$  into the wells of your well plate as shown by Figure 1.
2. Clean the copper, zinc, lead, and iron electrodes using steel wool. Rinse with deionized water. Place each metal electrode in its corresponding ionic solution (It is important that the correct metal is in the correct solution, or your cell will not work properly).
3. Obtain small strips of filter paper, to be used as salt bridges. Wet each strip with 0.1 M  $\text{KNO}_3$  and insert one end into the  $\text{KNO}_3$  solution and the other end into the  $\text{Cu}(\text{NO}_3)_2$ . Repeat for each of the three remaining salt solutions. (Hint: Make sure the strip is entirely wet. Any dry patches will inhibit the flow of ions through the bridge.)

- Attach an alligator clip from the middle hole on the bottom of the voltmeter to one of the metal electrodes and attach a second clip from the right hole on the bottom of the voltmeter to a different metal electrode. If the voltmeter reads a negative voltage, reverse the hookup so that each clip is now attached to the other metal in the pair.
- Record the voltage of the electrochemical cell for the reaction occurring at each electrode on the Report Sheet Part A.
- Repeat for the remaining cells.

Figure 1: Arrangement of ionic solutions for measuring cell potentials



**Report Sheet Part A:** Transfer the measured  $E_{\text{cell}}$  to column 4 of Table 1.

Cell	Measured $E_{\text{cell}}$	Anode	Cathode
Cu-Zn	_____	_____	_____
Cu-Pb	_____	_____	_____
Cu-Fe	_____	_____	_____
Pb-Zn	_____	_____	_____
Pb-Fe	_____	_____	_____
Fe-Zn	_____	_____	_____

Which of the four metals is the best oxidizing agent? Explain.

Which is the most effective reducing agent? Explain.

What purpose does the moist filter paper serve?

Complete column 6 of Table 1. Compare the measured potential and the calculated potential and provide a valid explanation of any major disagreement.

**Table 1.**

Cell	Spontaneous Standard Cell Reaction	Standard Cell Potential	Measured Potential	Calculated Potential	% Error
Cu-Zn					
Cu-Pb					
Cu-Fe					
Pb-Zn					
Pb-Fe					
Fe-Zn					

## Part B: Electrolysis

**Read and/or review Chapter 19: Sections 10-11 in your textbook.**

Electrolysis takes place in an electrolytic cell. Electrolysis occurs when the passage of an electric current through a solution causes chemical reactions at the electrodes that would otherwise be nonspontaneous. The current from an outside source causes the nonspontaneous transfer of electrons.

When electrolysis occurs in an aqueous solution, the oxidation and reduction of water are always possible reactions. These reactions are easily identified by gas evolution and pH changes in the solution near the electrode.

**Reduction (cathode, -) half-reaction for water:  $2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$**

**Oxidation (anode, +) half-reaction for water:  $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$**

1. Weigh 2.0 g of solid KI, in a 200 mL beaker. Add 100 mL of deionized water and mix until the salt dissolves.
2. Pour the solution into the U-tube provided. Using pH paper, measure the pH of the solution at each electrode.
3. Attach the alligator clips from the dc power supply (9V battery) to the carbon electrodes.
4. Record any observations (i.e. gas evolution, color changes).
5. Allow the electrolysis to proceed for 10 minutes.
6. Check the final pH at both electrodes. Remove the electrodes and examine. Record any observations and the probable identity of any substance that might be present.

Report Part B:

Initial pH of KI solution: \_\_\_\_\_

Initial  $[\text{OH}^-]$  : \_\_\_\_\_

Initial time: \_\_\_\_\_

Observations occurring within the U-tube during the reaction (i.e. color changes, gas evolution etc.), you may draw a picture if that helps your description:

Electrode observations:

Final time: \_\_\_\_\_

Final pH: \_\_\_\_\_

Final  $[\text{OH}^-]$  : \_\_\_\_\_

What are the possible competing reactions that can occur at the anode? Based on your observations which reaction does occur?

What are the possible competing reactions that can occur at the cathode? Based on your observations which reaction takes place?

What is the overall reaction for the cell?

Calculate the average current that flowed through the electrolytic cell from the  $\Delta[\text{OH}^-]$ , the volume of the solution, and the time of electrolysis.

## Part C: Electroplating Pennies

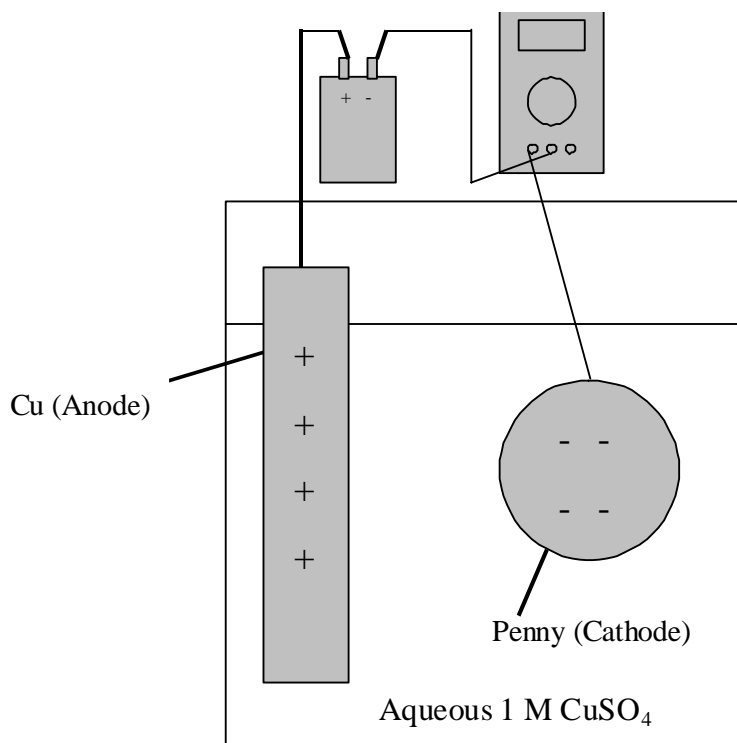
$1 \text{ C} = 1 \text{ A} \cdot \text{s}$  (C (coulomb), A (ampere), s (seconds))

Faraday Constant:  $1 \text{ F} = 96485 \text{ C/mole } e^-$

### Procedure:

1. Polish a penny and a strip of copper. Measure their mass ( $\pm 0.001 \text{ g}$ ). Set up the apparatus as shown in Figure 2, making sure to connect the copper strip to the positive terminal (anode) and the penny to the negative terminal (cathode) of the 9V battery.
2. Add 50 mL of 1.0 M  $\text{CuSO}_4$  to the the 100 mL beaker. Start timing the electrolysis.
3. Allow electrolysis to proceed for 15 min. Record the amperage periodically.
4. After 15 minutes remove the copper strip and the penny and allow to air dry. Record their masses.

Figure 2.



**Report Sheet (show all calculations):**

Record observations made of both the electrodes and the solution:

**Copper Strip:**

Initial Mass: \_\_\_\_\_ Final Mass: \_\_\_\_\_ Mass Change: \_\_\_\_\_

**Penny:**

Initial Mass: \_\_\_\_\_ Final Mass: \_\_\_\_\_ Mass Change: \_\_\_\_\_

Length of time of electrolysis: \_\_\_\_\_

Average Current (recorded in Amp): \_\_\_\_\_

Amount of Cu reduced (recorded in mol): \_\_\_\_\_

Convert mol of Cu to mol of electrons ( $e^-$ ): \_\_\_\_\_

Calculate the number of Coulombs that passed through the electrolytic cell: \_\_\_\_\_

Calculate the Faraday constant using your experimental information: \_\_\_\_\_

Compare your result with the literature value of the Faraday constant. % error: \_\_\_\_\_

**Show calculations:**